METHODS OF HEAT TRANSFER:

HEATING PROCESSES (heat released to surroundings)

- Freezing of liquid to solid.
- Condensation of vapor to liquid (reverse of boiling or evaporation).
- Compression of a gas.
- Condensation of vapor to solid (reverse of sublimation).

COOLING PROCESSES (heat absorbed from surroundings)

- Boiling of liquid to vapor.
- Evaporation of liquid to vapor.
- Melting of solid to liquid.
- Sublimation of solid to vapor.
- Expansion of a gas.
Chapter 18:
First Law of Thermodynamics: \( \Delta U = Q + W \). The change in internal energy of a system equals the heat added to the system plus the work done on the system.

A good test case is an ideal gas in a cylinder with a movable piston.

The work done ON the gas is

\[
W = - \int_{V_1}^{V_2} pdV,
\]

which is \(< 0\) if \( V_2 > V_1 \) (expansion), and \( > 0 \) if \( V_1 > V_2 \) (compression).

Types of Thermodynamic Processes:
- Isothermal, \( \Delta T = 0 \).
- Constant Volume, \( \Delta V = 0 \).
- Isobaric, \( \Delta p = 0 \).
- Adiabatic, \( Q = 0 \).

**Isothermal Process:**

\[
W = -NkT \ln \left[ \frac{V_2}{V_1} \right].
\]
For an ideal gas, $\Delta T = 0$ means $\Delta U = 0$.

**CONSTANT VOLUME PROCESS:**

$$\Delta U = Q$$

since no work is done on or by the system.

If $Q = nC_v\Delta T$ then $\Delta U = nC_v\Delta T$.

**ISOBARIC PROCESS:**

$$W = -p\Delta V, \quad Q = nC_v\Delta T + p\Delta V,$$

and using the Ideal Gas Law,

$$Q = n(C_v + R)\Delta T \equiv nC_p\Delta T.$$

**ADIABATIC PROCESS:**

$$\Delta U = W, \text{ since } Q = 0.$$

Using the Ideal Gas Law gives

$$\frac{dp}{p} + \frac{C_p}{C_v} \frac{dV}{V} = 0,$$

so that

$$pV^\gamma = \text{const.}, \text{ } TV^{\gamma^{-1}} = \text{const.}$$

where $\gamma = C_p/C_v$. 
**Equipartition of Energy:**

For an ideal gas, \( U = (3/2)NkT = (3/2)nRT \),

So for a monatomic gas, \( C_v = (3/2)R \).

For every additional degree of freedom, add \((1/2)kT\) to \( U \).

Thus for an atom in a solid, \( C_v = 3R \).

For a diatomic molecule at normal temperatures, \( C_v = (5/2)R \).
Chapter 19:

It is an obvious fact of nature that all readily-observable processes are **irreversible**. There is an **arrow of time**, flowing from past to future. Yet physicists quickly realized that it is by no means obvious where the irreversibility comes from. All fundamental laws of physics are **time-reversal invariant**, that is they have the same form if $t$ is replaced everywhere by $-t$. And the fundamental constants of nature are time-independent. So what is the origin of irreversibility of natural processes?

**Thermodynamic Entropy:** Clausius pointed out that the quantity

$$dS = \frac{dQ}{T}$$

irreversibly increases in thermodynamic processes. In the course of any physical process, $\Delta S > 0$ unless it is already at a maximum initially.

**Entropy:** Boltzmann pointed out that entropy is directly related to probability.

$$S = k \ln N_{ms}.$$  

Here $N_{ms}$ is the number of microstates that make up a given state of the system. So when you say
that "entropy increases," you are merely saying that a system subjected to accident tends to be left in a more probable state than the one it started in.

**The Second Law of Thermodynamics:**
In the course of a process, $\Delta S > 0$, in general.

**Heat Engine:**
Heat $Q_h$ is extracted from $T_h$ and a portion used to do work $W$. The remainder, $Q_c$, is expelled into an environment at $T_c$.

From energy conservation $Q_h = W + Q_c$. If we define the efficiency as $\epsilon = W/Q_h$, then we immediately find

$$\epsilon = 1 - \left[ \frac{Q_c}{Q_h} \right].$$

Because $\Delta S > 0$ in any cycle of the engine, it is impossible for $W = Q_h$ and $\epsilon = 1$.

**Carnot Efficiency:**
Physicists noticed that despite all engineering improvements the efficiency of all heat engines remained small. Carnot suggested comparing real engines to an imaginary engine for which $\Delta S = 0$. 
Then
\[ \Delta S = \frac{Q_c}{T_c} - \frac{Q_h}{T_h} = 0 \]
so that \( \frac{Q_h}{T_h} = \frac{Q_c}{T_c} \) and therefore
\[ \epsilon_c = 1 - \left[ \frac{T_c}{T_h} \right]. \]
Typically \( \epsilon \approx \epsilon_c / 2 \).
Different Statements of the 2nd Law:

- No process can convert 100% of a heat flow into work.

- In a closed system, any $\Delta T$ will eventually go to zero. ["Heat Death"]

- A system will wind up in more and more probable states as time goes by, unless it starts in a most probable state.

- All energy releases are eventually entirely dissipated as heat.

- All basic processes involving complex systems are irreversible. ["Time’s Arrow"] This does not necessarily apply to processes involving fundamental point particles.

Statistical Entropy $S = k_B \ln N_{ms}$.

Thermodynamic Entropy $\Delta S = Q/T$. 
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LEFT HALF

EXAMPLES

1 3 2 4

4

1

3

4

2

6

1 3 2 4

2

3 1

1 4

3 3

1 3 2 4

3 4 2 1

2 4 3

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Total = 16